

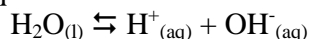
A Visual Introduction to Acid-Base Chemistry

Objectives:

1. Perform a neutralization reaction.
2. Perform reactions of acids and bases on various materials.
3. Test various solutions with acid-base indicators.
4. Measure the pH of various of common materials.
5. Perform dissociation reactions of salts and test the solutions pH.
6. Prepare buffer solutions and measure their buffer capacities.

Introduction:

In 1884 Svante Arrhenius proposed the first theoretical model for acids and bases. Prior to that time, these chemically opposite substances were described in properties such as their taste; their effects on metals, carbonates, and dyes (called indicators); their feel to the touch, and their ability to react with each other. According to the **Arrhenius theory**, pure water dissociates to some extent to produce hydrogen ions, H^+ and hydroxide ions, OH^- . When this occurs, equal amounts of H^+ and OH^- ions are produced:

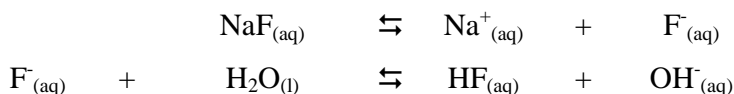


The list of **strong acids** include, HCl, HBr, HI, HNO_3 , H_2SO_4 and $HClO_4$. The list of strong bases include, LiOH, NaOH, KOH, RbOH and CsOH. What do the **strong bases** have in common? They are all composed of a group one metal and a hydroxide. From the solubility rules all compounds containing a group one metal are soluble. All of these acids and bases dissociate 100% in water.

pH is used to describe the acidic or basic nature of a solution. The H in pH means the concentration of hydrogen ion, H^+ in a solution. The p in pH is just a mathematical trick to make a decimal number a whole number. People do not like decimals and do not like exponents and really hate negative exponents. The p in pH is $-\log$. Again, the H in pH is the concentration of H^+ in molarity, but since it is in water we really mean hydronium, $[H_3O^+]$.

Salts are compounds composed of a metal and a non-metal. An interesting phenomenon occurs when certain salts dissociate in an aqueous solution, the pH may change. If the cation and anion of which this salt is composed are the cations and anions from strong acids and bases the pH of the solution does not change. NaCl is composed of the cation Na^+ and the anion Cl^- . If an H^+ is added to the Cl^- you get hydrochloric acid. Will this HCl molecule form in this aqueous environment? No, as HCl is a strong acid and strong acids dissociate 100% so you cannot have an HCl molecule in water. If an OH^- is added to the Na^+ from the salt will an NaOH molecule form? No, as NaOH is a strong base and strong bases dissociate 100%, so you cannot have an NaOH molecule in water.

What salts will the pH change? Salts of weak acid and weak bases. There are way too many weak acids and bases to list, if you remember all the strong acids and bases, then all the other millions of acid and bases in the world will be weak. So, an acid like HF, hydrofluoric acid is a weak acid. If a salt is made with the anion of that weak acid, that salt will be able to change the pH of a solution.

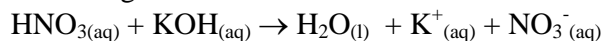


The above two equations show the dissociation of a salt and the basic properties of the anion of a weak acid. The fluoride ion is a base as it takes an H^+ ion from the water.

In general, the anions of weak acids behave as bases while the anions of strong acids do not. Similarly, the cations of weak bases behave as acids, while the cations of strong bases do not.

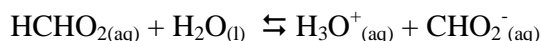
The term **neutralization** is often used to describe a reaction in which equal amounts of acid and base react with each other. Acid and bases can react with some compounds to change their color, these compounds are dyes that change color as the pH changes are referred to as indicators. This is one method used to determine the point at which exact amounts of acid and base have been reacted in neutralization reactions.

Normally, the products of the reaction are water and salt. The salt follows the rule of solubility and will either dissolve or not depending on what salt is formed. Thus, the neutralization of nitric acid with potassium hydroxide is represented by the following reaction:

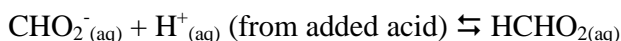


The potassium ion and nitrate ion are dissociated from each other as, all compounds containing a group one metal are always soluble. Also, all compounds containing nitrates are soluble. So, this compound is very soluble, neither ion wants to form a solid in water. The products of this reaction could be duplicated by simply adding the salt potassium nitrate to a container of water. To determine which salts will dissolve and which will remain as a solid molecule, consult a table of the rules of solubility.

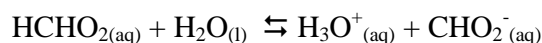
Solutions containing substances with the ability to donate protons substances with the ability to accept protons have the capacity to maintain a fairly constant pH despite the addition of acids or bases. Solutions with this capacity are called **buffers**, and the amount of acid or base they can accept without significant change is called the buffer capacity. A common type of buffer consists of a solution containing about equal amounts of a weak acid and the salt anion of the weak acid. Formic acid (HCHO_2) and its salt, sodium formate (NaCHO_2) could be used. Formic acid reacts with water as follows:



Thus HCHO_2 behaves as a Brønsted acid, while CHO_2^- would behave as a Brønsted base. In a solution containing both species, the addition of acid, H^+ would not change the pH much because the CHO_2^- ions would react as follows to minimize the formation of H_3O^+ :

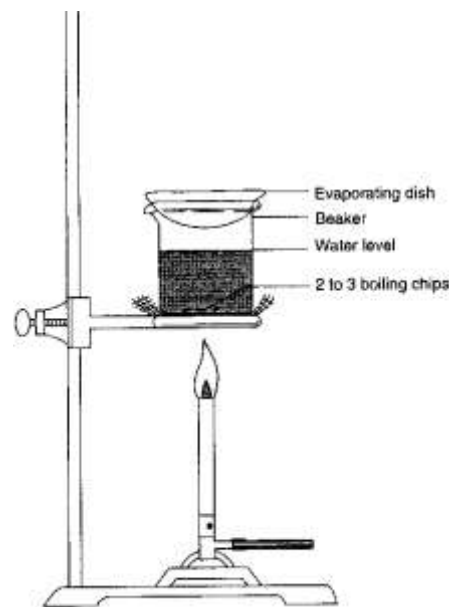


Similarly, any added base will deplete the H_3O^+ concentration, but the HCHO_2 molecules present will replenish them by undergoing the reaction from left to right in the below equation.



Procedure Part A:

1. Put 2.0 mL of 1.0 M sodium hydroxide solution (NaOH) into a test tube.
2. Using a glass stir rod, touch this solution to red and blue litmus paper. Record your results.
3. Put 2.0 mL of 1.0 M hydrochloric acid solution (HCl) into another test tube.
4. Using a glass stir rod, touch this solution to red and blue litmus paper. Record your results.
5. Pour the contents of one of the test tubes into the other test tube and mix the contents.



6. Pour the resulting solution into a clean evaporating dish and evaporate to dryness over a boiling water bath similar to the set up shown in the diagram.
7. While evaporation proceeds, move on to part B while you wait.
8. When evaporation is complete, collect the solid residue on a piece of filter paper, note and record it's appearance, and divide it into two equal portions.
9. Dissolve one portion in 20 drops of distilled water in a clean test tube.
10. Test the dissolved sample for chlorine ions by adding 2 drops of 0.1 M silver nitrate solution. A distinct clouding of the solution or the formation and settling of a white solid indicates the presence of Cl^- . Record the test results.
11. Moisten the end of a clean glass stirring rod with distilled water.
12. Dip the moist end of the rod into the second portion of collected solid so that solid particles adhere to the rod.
13. Place the end of the rod with adhering solid into the flame of your burner. The appearance of a bright orange-yellow flame confirms the presence of sodium ions. Record the results of this flame test.

Procedure Part B:

1. Use pH paper measure the pH of the following substances:
 - a. household ammonia
 - b. laundry detergent
 - c. dishwashing detergent
 - d. baking soda
 - e. aspirin
 - f. buffered aspirin
 - g. tap water
2. Use 2 different types of pH paper one wide range and one narrow range. Dip a clean glass stirring rod into the liquid being tested, and then touch the rod to each piece of pH paper. Compare the color of the spot on the paper to the chart that is located on the paper dispenser. **Thoroughly, clean the glass stirring rod between each test.**
3. If the samples listed above are solids, dissolve about 0.1 g in 2mL of deionized water and test the resulting solutions.
4. Record the measured pH values and tape the strips to your record sheet.

Procedure Part C:

1. On two separate glass plates, place one drop of each of the first 5 indicators.
2. On two other glass plates, place one drop of each of the second 5 indicators.
3. Place a drop of 0.05 M hydrochloric acid into each type of indicator.
4. Place a drop of 0.05 M sodium hydroxide into each type of indicator.
5. Mix well using a toothpick and record the resulting solution color.
 - a. phenolphthalein
 - b. methyl orange
 - c. methyl red
 - d. methyl blue
 - e. universal indicator pH 4-10
 - f. universal indicator pH 1-7

Procedure Part D:

1. Prepare each of the following solutions in separate, clean test tubes. Mark the tubes so you can remember their contents.
 - Solution A: 20 drops of 0.1 M sodium acetate solution ($\text{NaC}_2\text{H}_3\text{O}_2$)
plus 20 drops of 0.1 M acetic acid solution ($\text{HC}_2\text{H}_3\text{O}_2$)
 - Solution B: 20 drops of 0.1 M aqueous ammonia solution (NH_3) plus
20 drops of 0.1 M ammonium chloride solution (NH_4Cl)
 - Solution C: 20 drops of 0.1 M sodium acetate solution ($\text{NaC}_2\text{H}_3\text{O}_2$)
plus 20 drops of deionized water
 - Solution D: 20 drops of 0.1 M sodium chloride solution (NaCl) plus
20 drops of deionized water
2. Mix each solution well, then pour half of it into another clean test tube so you have two equal samples.
3. Use a pH paper, measure the pH of one sample of each solution. Record this initial pH value.
4. Also measure and record the pH of a sample of deionized water.
5. Add 8 drops of 0.05 M hydrochloric acid to one sample of each solution and to a 20-drop sample of deionized water. Mix well, measure, and record the resulting pH of each solution.
6. Repeat Step 5, using the other solution samples and a fresh sample of deionized water, but add 0.05 M sodium hydroxide solution instead of HCl.

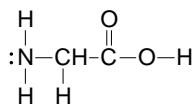
A Visual Introduction to Acid-Base Chemistry Student Handout

Pre-Lab Acid-Base

1. Define a(n):
 - a. acid:
 - b. base:
 - c. salt:
 - d. buffer:
2. What does pH tell you about a solution?
3. How do you calculate pH?
4. What is the difference between a strong acid and a weak acid?
5. Make a prediction to what would happen if you used pH paper with a range of 1-2.5 to test a solution with a pH of 8.

Concluding Questions:

1. At the end of the first part of this lab, what, specifically, is left in the evaporating dish?
2. What would be most effective in removing hard-water deposits, containing calcium carbonate, and other carbonates, from a sink, vinegar or household ammonia? Defend your choice with results from this lab.
3. Most cleaning agents (cleansers, soaps, and the like) tested in this lab are best classified as which of the following: neutral, acidic, or basic?
4. Which of the following reactions best represents the reaction of a salt that lowers pH when dissolved in water? Let C^+ and A^- represent the cation and anion of the salt.
 - a. $CA + H_2O \rightarrow C^+ + A^- + H_2O$
 - b. $CA + H_2O \rightarrow HA + C^+ + OH^-$
 - c. $CA + H_2O \rightarrow COH + A^- + H^+$
 - d. $CA + H_2O \rightarrow COH + HA$
5. Using a sample of a buffer solution, how could you determine buffering capacity?



6. Amino acids, like glycine shown here, can behave as buffers. Write out a chemical reaction showing glycine acting a buffer against an acid, then again against a base. Explain your answers.